INTERMOLECULAR FORCES

when predicting the shape of molecules (and poly atomic ions) we use:

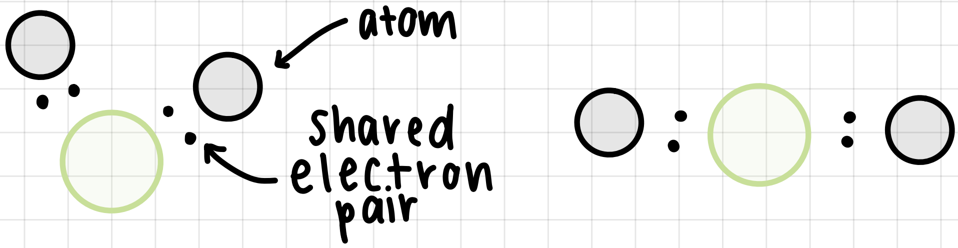
VSEPR

Valence Shell Electron Pair Repulsion Theory

& this means that:

valence electrons in pairs will repel each other and ∴ they will be

arranged as far away from each other as possible



lone pairs also repel other electrons

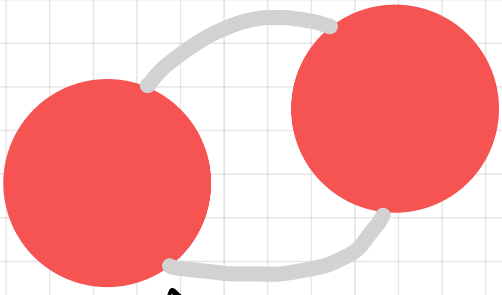
the repulsion of lone pairs > repulsion of shared electrons

∴ they push other electrons away more than shared electrons

molecules with 2 atoms (diatomic molecules)

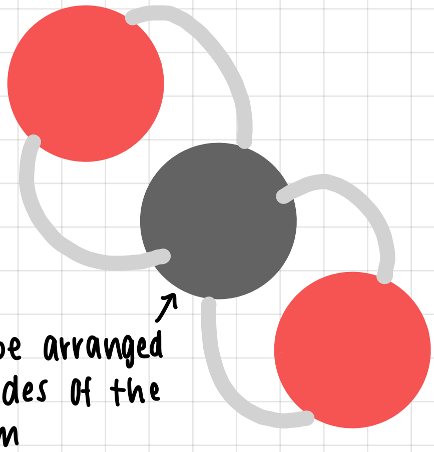
the shared electrons lie between the 2 molecules

LINEAR SHAPE



atoms will be in a straight line

molecules with 2 groups of electrons around the central atom



ATOMS ARE ARRANGED 180° APART

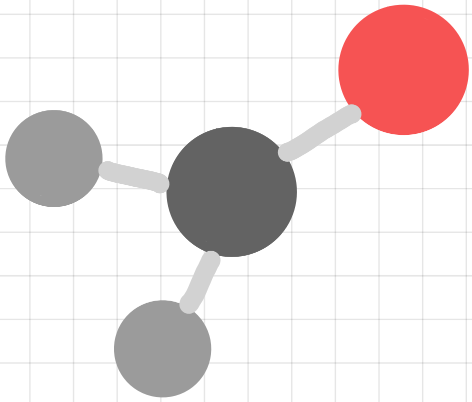
LINEAR SHAPE

atoms will be arranged on opposite sides of the central atom

molecules with 3 groups of electrons around the central atom

ATOMS ARE ARRANGED 120° APART

TRIGONAL PLANAR SHAPE



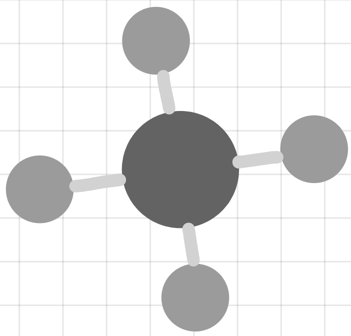
molecules with 4 groups of electrons around the central atom

GROUPS ARE ARRANGED 109·5° APART

the shape of the molecule depends on the number of bonding pairs and lone pairs

all electron pairs are bonding pairs:

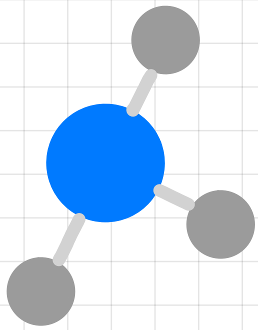
= no lone pairs



TETRAHEDRAL SHAPE

3 electron pairs are bonding pairs

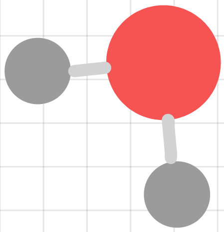
= 1 lone pair



PYRAMIDAL SHAPE

2 electron pairs are bonding pairs

= 2 lone pairs



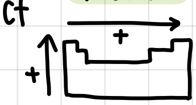
BENT TRIATOMIC SHAPE

FLOWCHART

ELECTRONEGATIVITY + POLARITY

ELECTRONEGATIVITY = the ability of an atom to attract electrons in a covalent bond

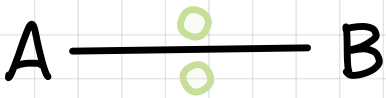
trend:



atoms with same/similar electronegativity

like O_2

like Hydrogen and Oxygen



- will share bonding electrons equally

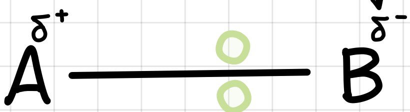
= 'pure' non-polar covalent bond

atoms with a difference in electronegativity

- difference between 0·5 and 1·8
  ↳ form covalent bonds but <u>don't</u> share the electrons equally

= polar covalent bonds

slight negative charge (delta negative)



electrons spend more time closer to the more electronegative element

examples:

F_2 = non-polar
:F - F: ← same electronegativity

HF = polar ← more electronegative
H - F:

the separation of charges is called a:

BOND DIPOLE

which can be drawn using a:

NET DIPOLE

e.g OH
3·44 2·20
= 1·24

$^{\delta-}O - H^{\delta+}$

direction points to negative side

net dipole

POLARITY = the measure of how polar a molecule or bond is

- the greater the difference in +ive and −ive ends of the dipole, the more polar it is

polarity can be determined by:

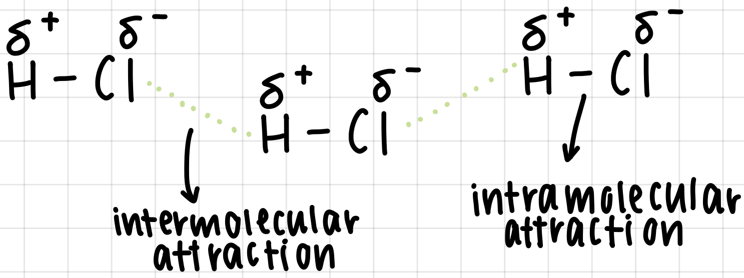
for diatomic molecules	for molecules w/ more than 2 atoms
↓ the difference in electronegativity	↓ - polarity of the covalent bonds - the shape of the molecule

* generally symmetrical molecules are non-polar

INTERMOLECULAR + INTRAMOLECULAR

INTRAMOLECULAR = forces that hold atoms within a molecule/compound (i.e. ionic, covalent & metallic bonds)

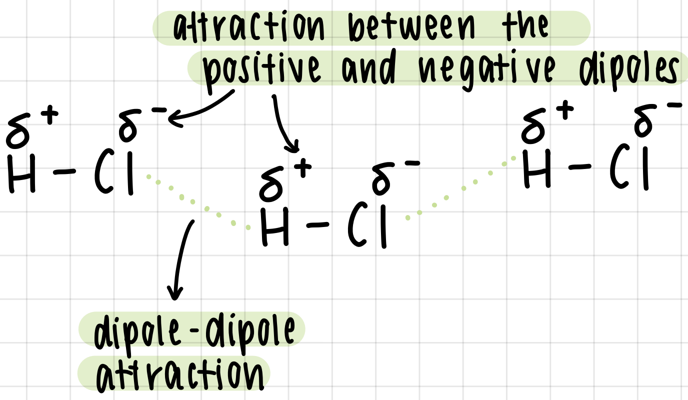
INTERMOLECULAR = forces that exist between covalent molecules



types of intermolecular forces

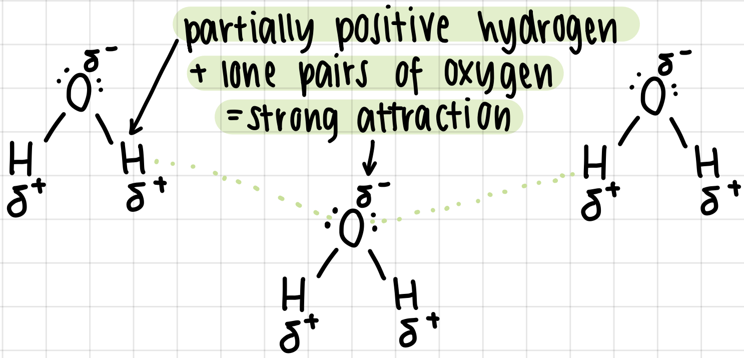
dipole - dipole force

- only in polar molecules
- relatively weak
  - ↳ because the partial charges in a dipole are small
- the more polar the molecule ⟶ stronger the dipole-dipole forces

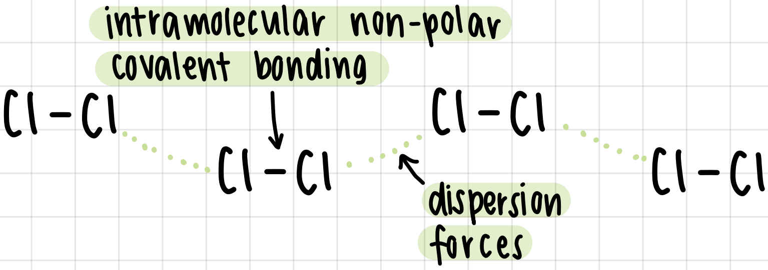


hydrogen bonding

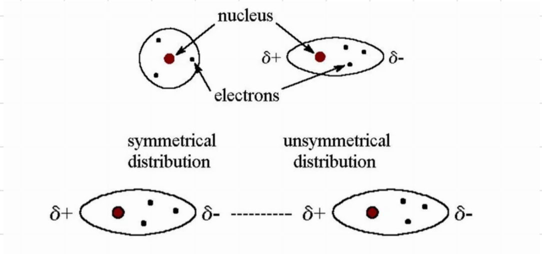
- oxygen, nitrogen and fluorine are all highly electronegative
  - ↳ ∴ they form very polar bonds
  - ↳ hydrogen has a very large partial positive charge (δ^+)
- relatively strong



dispersion forces



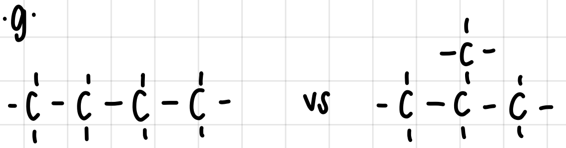
- weakest type of intermolecular force
- all molecules have dispersion forces
  - ↳ this explains the existence of intermolecular forces in non-polar substances despite the lack of a dipole
- occur due to movement of electrons
- at a particular moment, there will be more electrons at one end of a molecule
  - ↳ this creates a temporary (instantaneous) dipole
  - ↳ the temporary dipole induces a temporary dipole on an adjacent molecule
  - ↳ the opposite charges then attract each other



Temporary dipoles caused by the movement of electrons in a molecule.

- larger molecules have more electrons ∴ stronger dispersion forces (+ vice versa)
- molecules w/ longer chains have stronger dispersion forces because of their large surface area

e.g.



butane vs methylpropane

strength of intermolecular forces IS GENERALLY:

$\boxed{\text{dispersion} < \text{dipole-dipole} < \text{hydrogen bonds}}$

* may vary depending on the number of electrons + the shape

physical properties + intermolecular forces

VAPOUR PRESSURE = measure of the tendency for a substance to evaporate

- measured by determining the pressure exerted in kPa
- stronger intermolecular forces = lower vapour pressure

molecules are held strongly = less tendency to evaporate + less pressure exerted

as temp. increases, vapour pressure increases (because kinetic energy increases + molecules can escape their intermolecular forces easier)

MELTING + BOILING POINT

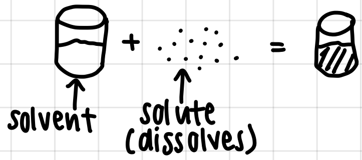
- stronger the intermolecular forces, the more energy needed to overcome the forces = higher melting + boiling point
- dispersion forces have their own trends in melting and boiling point (because of size and shape of molecules)

BOILING POINT

= temp. where vapour pressure = atmospheric pressure (the temp. where liquid turns to vapour)

SOLUBILITY

= the ability of a substance to dissolve



- covalent molecules don't break their bonds when dissolved (don't break into individual atoms)
- the intermolecular forces are disrupted which separates the covalent molecules making up the substance